

CH. 3 Practice

1. An element, E, has only two naturally occurring isotopes. E-10 has a mass of 10.01 amu and a natural abundance of 19.78%. E-11 has a mass of 11.01 amu and a natural abundance of 80.22%. What is the average atomic mass of E?

$$(10.01)(.1978) = 1.980$$

$$(11.01)(.8022) = 8.832$$

$$10.812 \text{ amu}$$

2. Chlorine has two stable isotopes. The mass of one isotope is 34.97 amu. Its relative abundance is 75.53%. What is the mass of the other stable isotope?

$$[(34.97)(.7553)] + [(x)(.2447)] = 35.45$$

$$[(x)(.2447)] = 35.45 - 26.41$$

$$[(x)(.2447)] = 9.04$$

$$x = 36.94 \text{ amu}$$

3. How many atoms are in 2.4 moles of neon gas? How many grams?

$$\frac{2.4 \text{ mol Ne}}{1} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol Ne}} = 1.4 \times 10^{24} \text{ atoms Ne}$$

4. How many grams of zinc are in 1.16×10^{22} atoms of zinc?

$$\frac{1.16 \times 10^{22} \text{ atoms Zn}}{1} \times \frac{1 \text{ mole Zn}}{6.022 \times 10^{23} \text{ atoms Zn}} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 1.26 \text{ g Zn}$$

5. How many mg of Br₂ are 4.6×10^{20} molecules of bromine?

$$\frac{4.6 \times 10^{20} \text{ molecules Br}_2}{1} \times \frac{1 \text{ mol Br}_2}{6.022 \times 10^{23} \text{ molecules Br}_2} \times \frac{159.80 \text{ g Br}_2}{1 \text{ mol Br}_2} \times \frac{1000 \text{ mg}}{1 \text{ g Br}_2} = 120 \text{ mg Br}_2$$

6.

How many grams are there in 0.36 moles of cobalt (III) acetate, $\text{Co}(\text{C}_2\text{H}_3\text{O}_2)_3$? How many grams of cobalt are in the sample? How many atoms of cobalt?

0.36 moles

$$\frac{0.36 \text{ moles } \text{Co}(\text{C}_2\text{H}_3\text{O}_2)_3}{1} \times \frac{236.08 \text{ g}}{1 \text{ mol } \text{Co}(\text{C}_2\text{H}_3\text{O}_2)_3} = 85 \text{ g } \text{Co}(\text{C}_2\text{H}_3\text{O}_2)_3$$

$$\frac{0.36 \text{ mol } \text{Co}}{1} \times \frac{58.93 \text{ g}}{1 \text{ mol } \text{Co}} = 21 \text{ g } \text{Co}$$

$$\frac{0.36 \text{ mol } \text{Co}}{1} \times \frac{6.022 \times 10^{23} \text{ atoms } \text{Co}}{1 \text{ mol } \text{Co}} = 2.2 \times 10^{23} \text{ atoms}$$

7.

Calculate the mass percent of chlorine in each of the following compounds.

A. ClF B. CuCl_2

$$\frac{35.45}{54.45} \times 100 = 65.11\% \text{ Cl}$$

$$\frac{70.90}{134.45} \times 100 = 52.73\% \text{ Cl}$$

8.

Chlorophyll a is essential for photosynthesis. It contains 2.72% magnesium by mass. What is the molar mass of chlorophyll a assuming there is one atom of magnesium in every molecule of chlorophyll a?

$$\frac{\text{g chl a}}{\text{mole chl a}} ?$$

$$\left(\frac{2.72 \text{ g Mg}}{100 \text{ g chl a}} \right)$$

$$\frac{1 \text{ mol Mg}}{1 \text{ mol chl a}} \times \frac{24.30 \text{ g Mg}}{1 \text{ mol Mg}} \times \frac{100 \text{ g chl a}}{2.72 \text{ g Mg}} = 893 \text{ g/mol}$$

$$\left(\frac{1 \text{ atom Mg}}{1 \text{ molec. chl a}} \right)$$

↓ so...

$$\left(\frac{1 \text{ mole Mg}}{1 \text{ mole chl a}} \right)$$

9.

Which of the following formulas can be empirical? Circle them.

A. CH_4 F. NH_4Cl B. CH_2 G. Sb_2S_3 C. KMnO_4 H. N_2O D. N_2O_5 I. CH_2O E. B_2H_8 → BH_4 lowest whole #
ratio of atoms

10.

A compound is found to contain 49.67% carbon, 48.92% chlorine, and 1.39% hydrogen. The molar mass of the compound is 289.9 g/mole. Determine the empirical and molecular formulas of the compound.

$$\frac{49.67 \text{ g C}}{12.01 \text{ g}} \times \frac{1 \text{ mol C}}{1} = 4.136 \text{ moles} = 3$$

$$\frac{48.92 \text{ g Cl}}{35.45 \text{ g}} \times \frac{1 \text{ mol Cl}}{1} = 1.380 \text{ moles} = 1$$

$$\frac{1.39 \text{ g H}}{1.01 \text{ g}} \times \frac{1 \text{ mol H}}{1} = 1.376 \text{ moles} = 1$$

Emp. formula: C_3ClH

Emp. mass 72.49 g

$$289.9 \div 72.49 = 4$$

molecular
formula: $\text{C}_{12}\text{Cl}_4\text{H}_4$

p119-120

69
71
73
75
77

11. Calcium carbonate decomposes upon heating, producing calcium oxide and carbon dioxide.

A. Write a balanced chemical equation for this reaction.



B. How many grams of calcium oxide will be produced after 12.25 grams of calcium carbonate are completely decomposed?

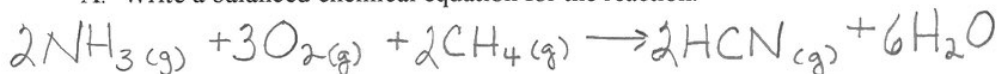
$$\frac{12.25 \text{ g CaCO}_3}{1} \times \frac{1 \text{ mole CaCO}_3}{100.09 \text{ g CaCO}_3} \times \frac{1 \text{ mol CaO}}{1 \text{ mol CaCO}_3} \times \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} = 6.864 \text{ g CaO}$$

C. What is the volume of carbon dioxide gas produced when the 12.25 grams of calcium carbonate completely decompose at STP?

$$\frac{12.25 \text{ g CaCO}_3}{1} \times \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = 2.74 \text{ L CO}_2$$

12. When ammonia gas, oxygen gas and methane (CH₄) gas are combined, the products are hydrogen cyanide gas and water.

A. Write a balanced chemical equation for the reaction.



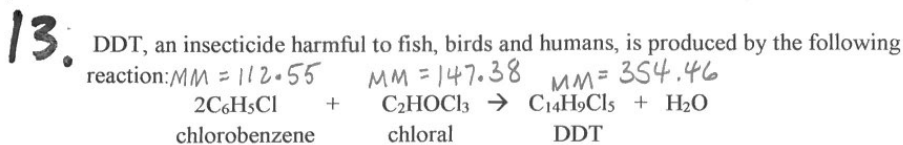
B. Calculate the mass of each product produced when 225 grams of oxygen gas is reacted with an excess of the other two reactants.

$$\frac{225 \text{ g O}_2}{1} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol HCN}}{3 \text{ mol O}_2} \times \frac{27.03 \text{ g HCN}}{1 \text{ mol HCN}} = 127 \text{ g HCN}$$

$$\frac{225 \text{ g O}_2}{1} \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 253 \text{ g H}_2\text{O}$$

C. If the actual yield of the experiment in (B) is 105 grams of HCN, calculate the percent yield.

$$\% \text{ Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{105}{127} \times 100 = 82.7 \% \text{ yield}$$



In a government lab, 1142 grams of chlorobenzene is reacted with 485 grams of chloral.

A. What mass of DDT is formed?

$$\frac{1142 \text{ g C}_6\text{H}_5\text{Cl}}{1} \times \frac{1 \text{ mol C}_6\text{H}_5\text{Cl}}{112.55 \text{ g C}_6\text{H}_5\text{Cl}} \times \frac{1 \text{ mol DDT}}{2 \text{ mol C}_6\text{H}_5\text{Cl}} \times \frac{354.46 \text{ g DDT}}{1 \text{ mol DDT}} = 1798 \text{ g DDT}$$

$$\frac{485 \text{ g chloral}}{1} \times \frac{1 \text{ mol chloral}}{147.38 \text{ g chloral}} \times \frac{1 \text{ mol DDT}}{1 \text{ mol chloral}} \times \frac{354.46 \text{ g DDT}}{1 \text{ mol DDT}} = 1170 \text{ g DDT}$$

Theoretical yield

B. Which reactant is limiting? Which is excess?



C. What mass of the excess reactant is left over?

$$\frac{485 \text{ g C}_2\text{HOCl}_3}{1} \times \frac{1 \text{ mol C}_2\text{HOCl}_3}{147.38 \text{ g C}_2\text{HOCl}_3} \times \frac{2 \text{ mol C}_6\text{H}_5\text{Cl}}{1 \text{ mol C}_2\text{HOCl}_3} \times \frac{112.55 \text{ g C}_6\text{H}_5\text{Cl}}{1 \text{ mol C}_6\text{H}_5\text{Cl}} = 741 \text{ g C}_6\text{H}_5\text{Cl used}$$

$$1142 \text{ g available} - 741 \text{ g used} = 401 \text{ g excess}$$

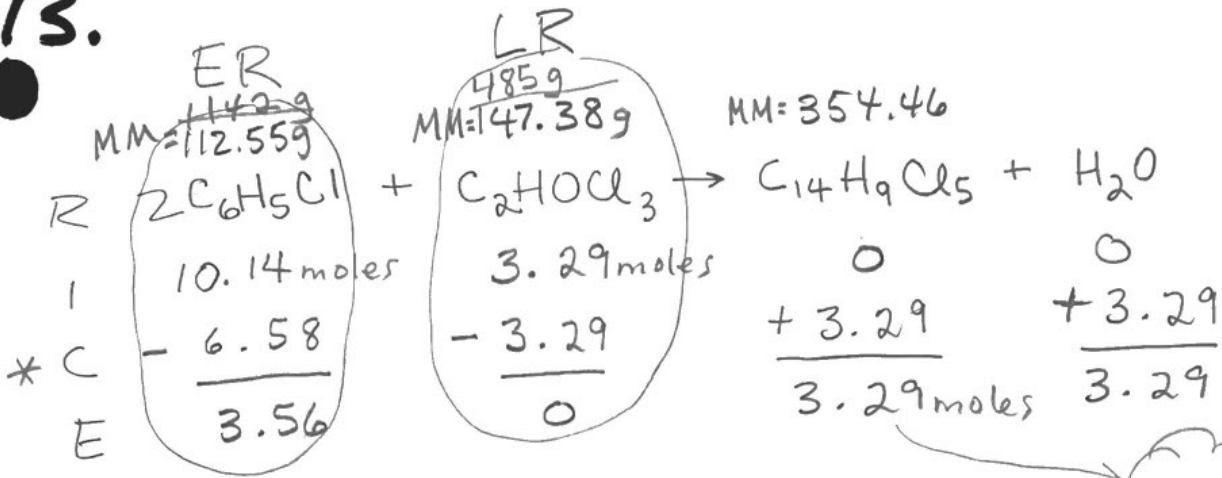
D. If the actual yield of DDT is 200.0 grams, what is the percent yield?

$$\frac{200.0 \text{ g}}{1166 \text{ g}} \times 100 = 17.15\% \text{ yield}$$

↓
RICE
chart
solution

RICE → moles or Pressure

13.



A) 1166 → 1170 g C₁₄H₉Cl₅ → TY

C) 401 g C₆H₅Cl excess

$$\% \text{Yield} = \frac{\text{AY}}{\text{TY}} \times 100$$

$$= \frac{200.0\text{g}}{1170\text{g}} \times 100 = 17.1\%$$